

The organic solution was dried over anhydrous magnesium sulfate, filtered, and evaporated to dryness. The residue was redissolved in 150 ml of ether and stirred overnight with 20 ml of 2 *N* aqueous hydrochloric acid. Usually, the oxindole had precipitated and could be collected by filtration. In case no precipitate was formed, the two layers were separated and the ethereal solution was concentrated causing the oxindole to crystallize. This latter procedure also allowed the isolation of a second fraction in those cases where the oxindole had precipitated initially.

**5-Methoxy-3-methylthio-2-oxindole (16a).** The oxindole, **16a**, was synthesized from 4-methoxyaniline (**1a**) *via* the above procedure. This allowed the isolation of 6.16 g of crude material that was recrystallized from methanol to give 4.90 g (0.023 mol, 53%) of pure 5-methoxy-3-methylthio-2-oxindole, mp 149.0–150.5°;  $\nu$  (KBr) 3400 and 3100 (NH), 1670  $\text{cm}^{-1}$  (C=O); pmr (DMSO- $d_6$ )  $\tau$  -0.34 (1 H, br s, NH), 3.00–3.22 (3 H, m, aryl H), 5.50 (1 H, s, CH), 6.26 (3 H, s, OCH<sub>3</sub>), and 8.00 (3 H, s, SCH<sub>3</sub>).

*Anal.* Calcd for C<sub>10</sub>H<sub>11</sub>NO<sub>2</sub>S: C, 57.40; H, 5.30; N, 6.69; S, 15.32. Found: C, 57.28; H, 5.36; N, 6.71; S, 15.20.

**5-Methoxy-2-oxindole (17a).** Desulfurization of 1.5 g (7.17 mmol) of 5-methoxy-3-methylthio-2-oxindole (**16a**) dissolved in 100 ml of absolute ethanol with W-2 Raney nickel gave, in 71% yield, 5-methoxy-2-oxindole, mp 148.5–150.5° (lit.<sup>18</sup> mp 152–154°).

**3-Methylthio-2-oxindole (16c).** 3-Methylthio-2-oxindole was synthesized from aniline (**1c**) *via* the above procedure. In this case the ethereal solution, containing the oxindole, was worked up by extracting the reaction mixture twice with 2 *N* aqueous hydrochloric acid, drying the organic phase over anhydrous magnesium sulfate, filtration, and evaporation. The solid residue was dissolved in 30 ml of refluxing ethanol, which was poured into 125 ml

of hot water. On standing, 5.10 g (0.029 mol, 65%) of 3-methylthio-2-oxindole crystallized and was collected by filtration, mp 126–127° (lit.<sup>5</sup> mp 126–127°).

**7-Methyl-3-methylthio-2-oxindole (16h).** The oxindole, **16h**, was synthesized from 2-methylaniline (**1h**) *via* the above procedure. There was obtained 5.30 g (0.028 mol, 62%) of **16h**, mp 190.5–193.0° (lit.<sup>5</sup> mp 194.0–195.5°).

**3-Methylthio-5-nitro-2-oxindole (16f).** The oxindole, **16f**, was synthesized from 4-nitroaniline (**1f**) *via* essentially the above procedure with the following modifications. A mechanically stirred solution of 0.044 mol of the chlorine-sulfide complex was prepared in the usual manner in 420 ml of methylene chloride. Through an addition funnel was added, as fast as possible, a solution of 13.1 g (0.088 mol) of 4-nitroaniline in a solution of 200 ml of methylene chloride and 25 ml of acetonitrile (to prevent premature crystallization of the nitro compound) with an accompanying temperature rise to -40°. The mixture was stirred for an additional 5 hr at -70° before the triethylamine was added. After warming to room temperature, a 200-ml portion of water was added and the layers were separated. Cyclization to the oxindole was accomplished with acid as described above, after which the layers were separated and the organic layer was extracted thoroughly with 2 *N* aqueous hydrochloric acid. After washing with saturated sodium bicarbonate solution, the organic phase was dried over anhydrous magnesium sulfate, filtered, and concentrated, leading to the crystallization of the product. There was obtained 1.15 g (0.5 mmol, 12%) of 3-methylthio-5-nitro-2-oxindole, mp 190.0–193.5° (lit.<sup>5</sup> mp 196–197°).

**Acknowledgment.** We are indebted to the National Cancer Institute of the Public Health Service for a grant which partially supported this investigation.

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## Oxidation of Hydrocarbons. IV. Kinetics and Mechanism of the Oxidative Cleavage of Cinnamic Acid by Acidic Permanganate<sup>1</sup>

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**Abstract:** Permanganate ion reacts rapidly with *trans*-cinnamic acid in aqueous perchloric acid to form an intermediate which subsequently decomposes in a slightly slower reaction to give benzaldehyde and Mn<sup>III</sup>. Inverse secondary deuterium kinetic isotope effects ( $k_D/k_H = 1.3$ ) which are obtained for the oxidation of *trans*-cinnamic acid- $\alpha$ -*d* and *trans*-cinnamic acid- $\beta$ -*d* are taken as evidence that the initial reaction involves addition of permanganate ion to the carbon-carbon double bond to form an intermediate hypomanganate ester. Normal secondary deuterium kinetic isotope effects ( $k_H/k_D = 1.09$ ) are observed for the decomposition of the intermediate into products. The first step of the reaction (ester formation) is characterized by a small enthalpy of activation (4.2 kcal/mol) and a large negative entropy of activation (-30 eu), whereas the second step (ester decomposition) exhibits a larger enthalpy of activation (11.9 kcal/mol) and an entropy of activation of similar magnitude (-27 eu). The rate of formation of the intermediate is rather insensitive to the presence of substituents on the aromatic nucleus; however, the rate of decomposition of the intermediate exhibited a Hammett  $\rho$  value of -1.1. An attempt has been made to propose a unified mechanism which is consistent with all of the known physical data concerning the reaction and which also satisfactorily accounts for the array of products that can be obtained under various conditions.

The oxidation of carbon-carbon bonds by permanganate ion is an important and well-known reaction in organic chemistry. Under alkaline conditions olefins are converted to the corresponding diols in good yields,<sup>2,3</sup> while in neutral or only slightly basic solutions

$\alpha$ -hydroxy ketones are produced.<sup>3,4</sup> These reactions are always accompanied by a certain amount of carbon-carbon bond cleavage<sup>5</sup> and under acidic conditions cleavage products predominate.<sup>6,7</sup>

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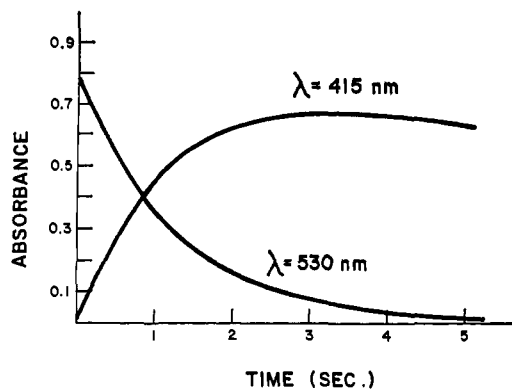


Figure 1. Reaction rates monitored by following the disappearance of the permanganate ion ( $\lambda$  530 nm) or the appearance of the intermediate hypomanganate ester ( $\lambda$  415 nm).

There is general agreement that cyclic manganese esters are formed in the initial step of this reaction.<sup>8a</sup> For example, it has been shown that during the dihydroxylation reaction the two hydroxy groups are added *cis*<sup>9</sup> and that the oxygens are transferred directly from the  $\text{MnO}_4^-$  ion.<sup>10</sup> More recently, species with transient lifetimes which are believed to be intermediate hypomanganate esters have been detected using stopped flow techniques.<sup>11,12</sup>

In the following paper we wish to describe, in detail, the results obtained from a study of the oxidation of *trans*-cinnamic acid in aqueous perchloric acid solutions. The use of *trans*-cinnamic acid as a substrate offers several distinct advantages. (i) The absorption spectrum of the intermediate is quite different from that of permanganate ion and may be readily detected either visually or spectrophotometrically.<sup>12</sup> (ii) The carboxylate group gives the molecule sufficient solubility in aqueous solutions. (iii) The presence of an aromatic ring permits a study of the effect of substitution on the rates of formation and decomposition of the intermediate. (iv) The vinyl hydrogens may readily be replaced by deuterium.<sup>13</sup>

### Experimental Section

**Reagents, Solvents, Solutions.** Baker analyzed reagent grade potassium permanganate was used as the oxidant in all experiments. The substituted *trans*-cinnamic acids were obtained from the Aldrich Chemical Co. and repeatedly recrystallized until they gave sharp melting points in agreement with literature values. Preparation of the deuterated compounds has previously been described.<sup>13</sup>

Distilled water, obtained from an all-glass still, was slowly redistilled from alkaline potassium permanganate before being used. Fisher analyzed reagent grade acetic acid was distilled under reduced pressure from acidic sodium dichromate.<sup>14</sup> The acidity of all solutions was adjusted by using Baker analyzed reagent grade perchloric acid.

A stock solution of approximately 0.025 *M*  $\text{KMnO}_4$  was prepared and stored at  $-10^\circ$ . Under these conditions the solution remained

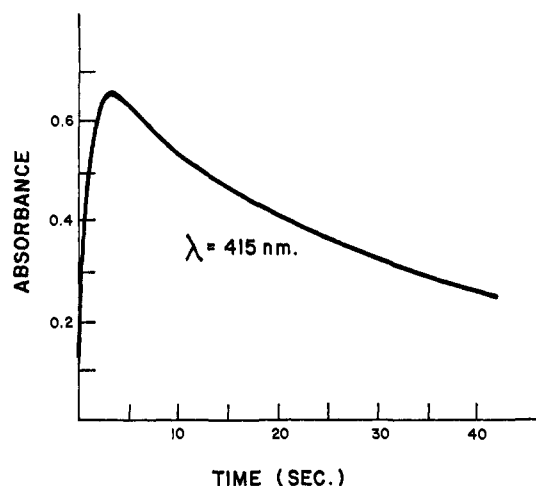


Figure 2. Plot showing changes in absorbance which take place as the intermediate forms and subsequently decomposes.

stable as indicated by the lack of any  $\text{MnO}_2$  deposited on the walls of the flask.

Stock solutions of aqueous *trans*-cinnamic acid were prepared by suspending the solid in an aqueous perchloric acid solution of known acidity and continuously shaking for several hours. Any undissolved acid was then removed by filtration and the concentration determined spectrophotometrically, using the absorption band at 270 nm.<sup>15</sup>

The solutions used in kinetic experiments were prepared by dilution of the stock solutions with aqueous perchloric acid of the desired concentration. Before being used in the stopped flow apparatus these solutions were degassed by freezing and then slowly thawing under reduced pressure ( $\sim 70$  mm). Since acidic permanganate solutions decompose slowly, the concentration of the degassed solutions of oxidant were determined spectrophotometrically at the start of each experiment.

Because of their limited solubility in water, it was necessary to use a mixed solvent system to study the rates of oxidation of the substituted cinnamic acids. Aqueous acetic acid (50% by volume) was found to be a suitable solvent. An appropriate amount of the substituted cinnamic acid was dissolved in 25 ml of purified acetic acid and diluted to 50 ml with 2.0 *M*  $\text{HClO}_4$  to give a reductant solution that was approximately  $1.0 \times 10^{-3}$  *M* in cinnamic acid, 50% by volume acetic acid, and 1.0 *M*  $\text{HClO}_4$ . The oxidant solutions were prepared in a similar fashion and were always about  $4.5 \times 10^{-4}$  *M*  $\text{KMnO}_4$ . Prior to use in kinetic experiments these solutions were also degassed as described above.

Because permanganate slowly attacks acetic acid under these conditions, it was necessary to make a small correction to all of the rates measured in this mixed solvent. Consequently the rates reported for the substituted cinnamic acids are slightly less reliable than the rest of the data reported in this paper.

**Kinetic Methods.** The formation of an intermediate in the oxidation of *trans*-cinnamic acid by aqueous acidic permanganate can easily be observed qualitatively. When a purple solution of the oxidant ( $2 \times 10^{-4}$  *M*) and a colorless solution of the reductant ( $5 \times 10^{-4}$  *M*) are mixed in a test tube, the resulting solution (light purple to begin with) becomes yellow within a few seconds and finally colorless after a few minutes.

Since the half-lives of these reactions are in the order of a few seconds, it was convenient to study them quantitatively using a Durrum Model D-110 stopped flow spectrophotometer. When the monochromator was set at the absorption maximum of permanganate ion (530 nm), the oscilloscope display clearly indicated a continuous decrease in absorbance as shown in Figure 1.

The ultraviolet spectrum of the intermediate was investigated by varying the wavelength setting and noting the absorbance about 3–4 sec after the reaction had been initiated. In this way an absorbance maximum was located at 415 nm. Since permanganate ion is transparent in this region, it was convenient to use this maximum to monitor the formation (Figure 1) and decomposition (Figure 2) of the reaction intermediate.

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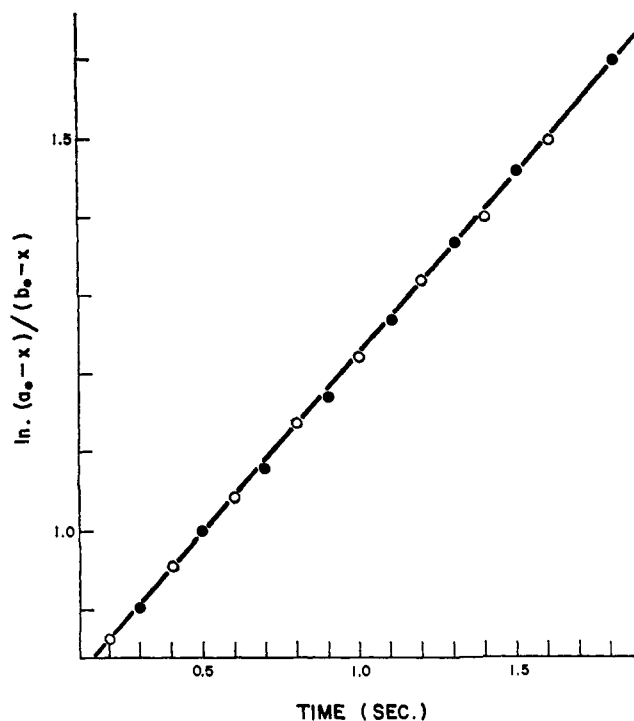
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**Figure 3.** Typical second-order rate plot. Points represented by open circles were obtained by monitoring the disappearance of permanganate ion at 530 nm. Points represented by closed circles were obtained by monitoring the appearance of the intermediate hypomanganate ester at 415 nm.

The kinetics of the initial reaction (formation of the intermediate) were investigated by spectrophotometrically monitoring the disappearance of the intermediate (at 415 nm). If in this reaction one molecule of permanganate ion combines with one molecule of *trans*-cinnamic acid to form the intermediate species, then second-order kinetics would be observed with the reaction being first order in both permanganate ion and cinnamic acid. If  $a_0$  and  $b_0$  are the initial concentrations of *trans*-cinnamic acid and permanganate ion, respectively, and  $x$  is the decrease in permanganate ion concentration at time  $t$  (and therefore also the decrease in cinnamic acid concentration), the rate of the reaction is described by eq 1.<sup>16</sup>

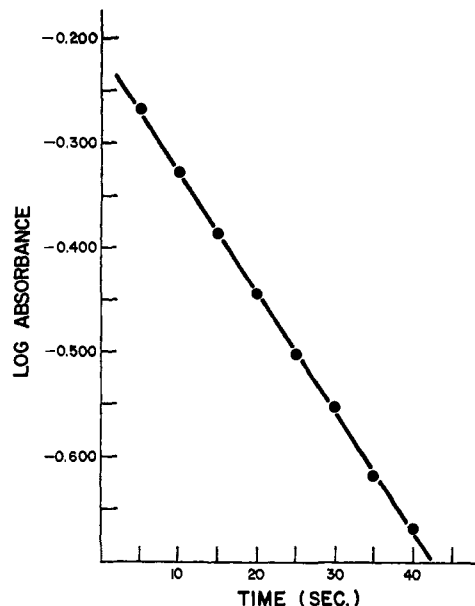
$$\ln (a_0 - x)/(b_0 - x) = (a_0 - b_0)k_2t + \ln a_0/b_0 \quad (1)$$

The corresponding plot for a typical rate determination has been reproduced in Figure 3. Similar linear plots, with high correlation coefficients, were always obtained thus indicating that the initial step of the reaction is second order. It should be noted that the points obtained by monitoring the appearance of the intermediate at 415 nm are directly superimposable on those obtained by following the disappearance of permanganate at 530 nm (Figure 3). This confirms that the disappearance of permanganate ion and the formation of the intermediate are the same process, at least insofar as it is currently possible to distinguish between them.

The concentrations of permanganate ion and cinnamic acid were carefully chosen so that the reductant was always in excess of (at least double) the concentration of the oxidant but not in so large an excess that it could be considered constant. Furthermore, the concentrations of both were *not* allowed to be approximately equal since the difference,  $a_0 - b_0$ , appears in eq 1 and any error in the concentrations would then result in a large error in the calculated rate constants.

Experimentally it was found more convenient to monitor the reaction at 530 nm, and the rate constants for the initial step ( $k_2$ ) were obtained by monitoring the disappearance of permanganate at this wavelength.

The order of the second detectable reaction (decomposition of the intermediate) could be determined by following the decrease in absorbance at 415 nm which is observed when the reaction is



**Figure 4.** Typical first-order plot of the decomposition of the intermediate hypomanganate ester.

studied over a longer time period (Figure 2). When the data were analyzed by use of the well-known integral rate laws<sup>16</sup> for zero-, first-, and second-order reactions, it was found that only first-order plots were linear. A typical example has been reproduced in Figure 4.

**Product Analysis.** Although it is known that benzaldehyde is the major organic product obtained when cinnamic acid is oxidized by permanganate under alkaline conditions,<sup>9</sup> it seemed advisable to also attempt to identify the products of the reaction under approximately the same experimental conditions that were used for the kinetic determinations.

*trans*-Cinnamic acid (1.13 g,  $7.64 \times 10^{-3}$  mol) and 250 ml of 1 M HClO<sub>4</sub> were added to a 500-ml round-bottomed flask fitted with a dropping funnel and a mechanical stirrer. The flask was flushed with nitrogen and suspended in a water bath maintained at a temperature of 30°. Then 0.60 g ( $3.82 \times 10^{-2}$  mol) of KMnO<sub>4</sub>, dissolved in 250 ml of 1 M HClO<sub>4</sub>, was added dropwise. The resulting solution, dark yellow in color, was allowed to stir for a few minutes after the last addition of potassium permanganate during which time the dark yellow color decreased in intensity. The solution was then extracted with  $3 \times 75$  ml of ether and the ether layer in turn extracted with dilute NaOH. The resulting ether layer was dried over anhydrous CaCl<sub>2</sub> and evaporated, a nitrogen atmosphere being maintained at all times. An infrared spectrum exhibited a carbonyl band at 1700 cm<sup>-1</sup> and certain other bands below 900 cm<sup>-1</sup> that are characteristic of benzaldehyde.<sup>17</sup>

The dilute NaOH solution was acidified and then extracted with  $3 \times 75$  ml of ether. The ether layer was dried over anhydrous CaCl<sub>2</sub> and evaporated to dryness leaving a white solid residue which was found to be cinnamic acid.

When an inert nitrogen atmosphere was not maintained, the only product obtained was benzoic acid. Thus it appears that benzaldehyde is the initial product but that it undergoes further oxidation to benzoic acid if oxygen is present.

Since the final solutions obtained from these reactions are colorless the manganese must be in either the +2 or the +3 oxidation state. Both manganese(II) and manganese(III) dissolve in water to give transparent solutions;<sup>8</sup> however, it is possible to distinguish between them because manganese(III) is known to disproportionate to manganese(II) and manganese(IV) under alkaline conditions.<sup>18</sup> The manganese(IV) species, MnO<sub>2</sub>, is insoluble and precipitates. When solutions remaining from kinetic determinations were made alkaline by addition of NaOH, a brown precipitate characteristic of MnO<sub>2</sub> was indeed obtained. Hence, it appears that under these conditions permanganate is reduced to manganese(III) while cinnamic acid is oxidized to benzaldehyde.

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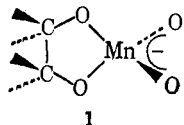
**Table I.** Rate Constants for the Oxidation of Substituted *trans*-Cinnamic Acids by Permanganate Ion

Substrate	Solvent	$\lambda_{\max},^a$ nm	$[\text{HClO}_4], M$	Temp, deg	$k_2, M^{-1} \text{sec}^{-1}{}^b$	$k_1, \text{sec}^{-1}{}^c$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O	415	0.99	30.0	$1.56 \times 10^3$	$2.60 \times 10^{-2}$
<i>trans</i> -Cinnamic acid- $\alpha$ - <i>d</i>	H <sub>2</sub> O		0.99	30.0	$2.02 \times 10^3$	$2.39 \times 10^{-2}$
<i>trans</i> -Cinnamic acid- $\beta$ - <i>d</i>	H <sub>2</sub> O		0.99	30.0	$2.08 \times 10^3$	$2.38 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		1.64	30.0	$1.94 \times 10^3$	$2.88 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		2.26	30.0	$2.14 \times 10^3$	$3.36 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		2.92	30.0	$2.36 \times 10^3$	$3.81 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		0.99	11.7	$9.65 \times 10^2$	$7.46 \times 10^{-3}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		0.99	21.1	$1.27 \times 10^3$	$1.19 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		0.99	30.2	$1.59 \times 10^3$	$2.63 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	H <sub>2</sub> O		0.99	41.1	$2.17 \times 10^3$	$5.59 \times 10^{-2}$
<i>trans</i> -Cinnamic acid	50% AcOH	393	1.01	30.0	$4.42 \times 10^2$	$1.92 \times 10^{-2}$
<i>p</i> -Methyl- <i>trans</i> -cinnamic acid	50% AcOH	403	1.01	30.0	$4.07 \times 10^2$	$1.99 \times 10^{-2}$
<i>p</i> -Chloro- <i>trans</i> -cinnamic acid	50% AcOH	405	1.01	30.0	$3.80 \times 10^2$	$7.28 \times 10^{-3}$
<i>m</i> -Chloro- <i>trans</i> -cinnamic acid	50% AcOH	417	1.01	30.0	$3.78 \times 10^2$	$4.78 \times 10^{-3}$
<i>p</i> -Nitro- <i>trans</i> -cinnamic acid	50% AcOH	400	1.01	30.0	$3.29 \times 10^2$	$2.45 \times 10^{-3}$

<sup>a</sup> Absorption maximum of the intermediate hypomanganate ester. <sup>b</sup> Second-order rate constants for the reaction between *trans*-cinnamic acid and  $\text{MnO}_4^-$  as defined by eq 3. <sup>c</sup> First-order rate constants for the disappearance of the intermediate ester as defined by eq 3. All values of rate constants are the average of three or more experiments with agreement being  $\pm 3\%$  or better.

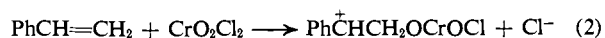
## Results and Discussion

As indicated by Figures 1–4 the reaction appears to consist of two consecutive steps, the first one being a bimolecular formation of an intermediate and the second a unimolecular decomposition of the intermediate. Since several previous workers<sup>10, 11, 19, 20</sup> have suggested that cyclic manganese(V) esters, **1**, were



likely intermediates in the reaction, it was of interest to test this possibility and to investigate solvent and substituent effects on the rates of both steps of the reaction.

As we have pointed out in a previous brief communication,<sup>12</sup> the results of deuterium substitution are entirely consistent with a mechanism in which an ester such as **1** is an intermediate in the oxidation of cinnamic acid. The rate of the initial bimolecular reaction is increased ( $k_D/k_H = 1.3$ ) by replacement of either the  $\alpha$  or the  $\beta$  protons by deuterium. This suggests that the hybridization of *both* carbon atoms is changing from  $sp^2$  to  $sp^3$ <sup>21</sup> and that the transition state is symmetrical with respect to the extent of rehybridization at both carbons. If attack had taken place at only one of the carbons before the transition state was achieved, it would have been expected to produce an inverse isotope effect only at that carbon. For example, the oxidation of styrene by  $\text{CrO}_2\text{Cl}_2$  exhibits an isotope effect only at the  $\beta$  carbon<sup>22</sup> suggesting that the rate-determining reaction produces a carbonium ion in which the  $\alpha$  carbon remains  $sp^2$  (eq 2). On the other hand,



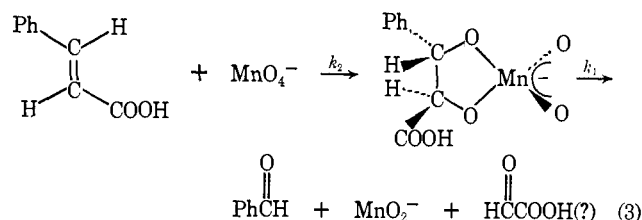
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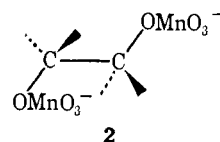
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the identical effects observed for the oxidation of both cinnamic acid- $\alpha$ -*d* and cinnamic acid- $\beta$ -*d* by permanganate are more consistent with the formation of a cyclic intermediate (eq 3). The only other possibility would



be formation of a diester (such as **2**), but this type of



intermediate is eliminated on the basis of kinetic evidence which indicates that the transition state for this reaction includes only one molecule of the permanganate ion.

The observation that the electronic absorption spectrum of the intermediate in aqueous perchloric acid solutions has a maximum at about 415 nm is not inconsistent with the suggestion that the intermediate may be some form of manganese(V).<sup>8c</sup> Moreover, the fact that the position of the maximum is also influenced by substituents on the cinnamic acid (Table I) substantiates that this intermediate is an ester and not just dissolved hypomanganate ions. If this intermediate manganese species was not combined in some way with the cinnamic acid moiety substituent, changes should have no effect on the spectrum.

When the rates at which the substituted cinnamic acids reacted were investigated, it was observed that the first step of the reaction was relatively insensitive to the nature of the substituents but that the second step was substantially retarded by electron-withdrawing substituents (Table I) with the Hammett  $\rho$  value being  $-1.1$  (Figure 5). Wiberg and Geer<sup>20</sup> have also noted

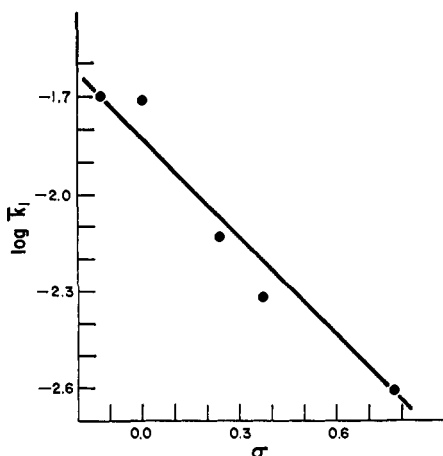


Figure 5. Hammett plot for the hypomanganate ester decomposition reaction.

that substituents have a small effect on the rate of disappearance of permanganate ion when cinnamic acids are oxidized in basic or neutral solutions.

The large effect of substituents on the ester decomposition reaction is probably associated with the development of a carbonyl in the second step of the reaction. Electron-donating substituents would increase the stability of the transition state if it possessed considerable carbonyl character. The normal secondary isotope effects observed for this step ( $k_H/k_D = 1.09$ ) are also consistent with the suggestion that the  $\alpha$  and  $\beta$  carbons have developed considerable carbonyl character by the time that the transition state for this step is realized.<sup>21</sup>

It is of interest to note at this point that the only intermediate detected by Simándi and Jáky<sup>6</sup> during the oxidation of fumaric acid by  $MnO_4^-$  was  $Mn^{3+}$ , although the second-order rate constants were similar to those which we have obtained. Hence, it appears that the rate of ester formation is about the same for cinnamic acid and fumaric acid under acidic conditions but that the rate of decomposition to manganese(III) is much faster in the case of fumaric acid. This suggestion is entirely consistent with the observation that the rate of ester decomposition, but *not* ester formation, is highly dependent on the presence of substituents on the aromatic nucleus of cinnamic acid.

Both the rate of formation and the rate of decomposition of the intermediate were found to be only slightly influenced by increases in the acidity of the medium (Table I). In fact, the second-order rate constants for reaction between permanganate ion and carbon-carbon multiple bonds appear to be rather insensitive to medium effects unless they cause a change in the formal charge on the reductants. Anions react more slowly than neutral molecules. For example, Wiberg and Geer<sup>20</sup> have reported a rate constant of about  $500 M^{-1} sec^{-1}$  for the oxidation of cinnamate ion by  $MnO_4^-$ , while we have found a value of approximately  $1300 M^{-1} sec^{-1}$  for the corresponding reaction with cinnamic acid. Simándi and Jáky<sup>6</sup> have observed very similar changes in rates for the oxidation of fumaric acid and its anions, as well as for acetylenedicarboxylic acid and its anions.<sup>23</sup>

From a study of the way in which the rates of these reactions respond to temperature changes it is possible

(23) L. I. Simándi and M. Jáky, *Tetrahedron Lett.*, 3489 (1970); *J. Chem. Soc., Perkin Trans. 2*, 1481 (1972).

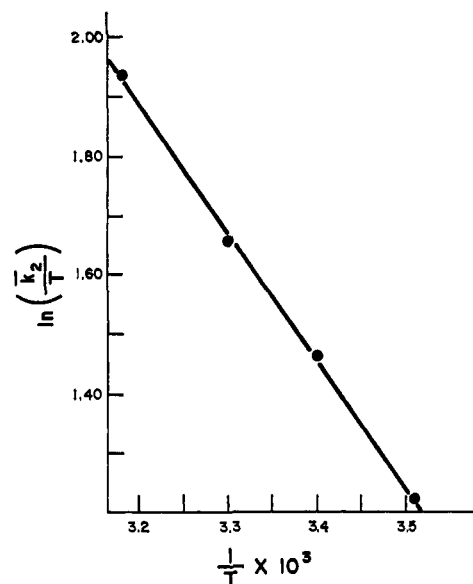


Figure 6. Plot for calculating the activation parameters for intermediate formation ( $\lambda$  530 nm):  $\Delta H^\ddagger = 4.2 \pm 0.5$  kcal/mol,  $\Delta S^\ddagger = -30 \pm 2$  eu,  $\Delta F^\ddagger = 13 \pm 1$  kcal/mol.

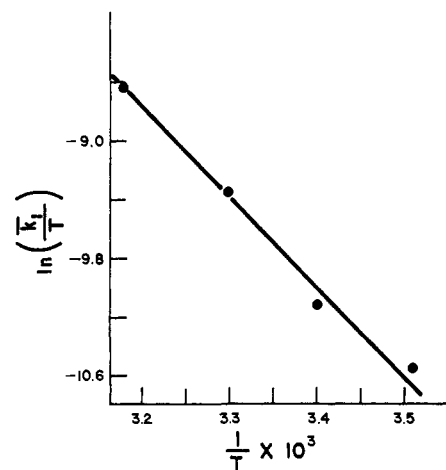


Figure 7. Plot for calculating the activation parameters for intermediate decomposition ( $\lambda$  415 nm):  $\Delta H^\ddagger = 11.9 \pm 0.5$  kcal/mol,  $\Delta S^\ddagger = -27 \pm 2$  eu,  $\Delta F^\ddagger = 20 \pm 1$  kcal/mol.

to determine the activation parameters,  $\Delta H^\ddagger$  and  $\Delta S^\ddagger$  (Figures 6 and 7). The ester forming reaction is characterized by a large negative entropy of activation ( $-30$  eu), which is certainly consistent with the formation of a cyclic intermediate in a bimolecular reaction. It is more surprising to observe that  $\Delta S^\ddagger$  for the second step is of approximately the same magnitude ( $-27$  eu). Since this step involves a decomposition of the cyclic intermediate into two or more parts, it might have been expected to exhibit a more favorable entropy of activation. Perhaps the explanation lies in what has been called the "external entropy," *i.e.*, entropy effects associated with solvation.<sup>24</sup> If the individual products were more highly solvated than the intermediate, an unfavorable entropy term would result.

The difference in the rates of the first and second steps is almost entirely attributable to differences in activation

(24) J. W. Larson and L. G. Hepler, "Solute-Solvent Interactions," J. F. Coetzee and C. D. Ritchie, Ed., Marcel Dekker, New York, N. Y., 1969.

enthalpies (4.2 kcal/mol for ester formation and 11.9 kcal/mol for ester decomposition). These values probably reflect a difference in the amount of bond breaking required in each step. For the first step cleavage of only one  $\pi$  bond is required, whereas the second step requires cleavage of three  $\sigma$  bonds, each step being accompanied by a lowering of the oxidation potential of manganese and formation of several new bonds.

The free energies of activation for the two successive steps are 13 and 20 kcal/mol clearly indicating that under these conditions the second step is rate limiting. We may note a remarkable similarity between these activation parameters and those previously reported (Table II). This emphasizes the probability that all of

**Table II.** Activation Parameters for the Reaction between Permanganate Ion and Carbon-Carbon Double and Triple Bonds

Reductant	Conditions	$\Delta H^\ddagger$ , kcal/mol	$\Delta S^\ddagger$ , eu	$\Delta F^\ddagger$ , kcal/ mol
Fumarate ion <sup>a</sup>	pH $\geq 4.9$	3.1	-34	13
Cinnamate ion <sup>b</sup>	pH 6-13	3.3	-36	14
Cinnamic acid <sup>c</sup>	0.99 M HClO <sub>4</sub>	4.2	-30	13
Acetylene- dicarboxylate ion <sup>d</sup>	pH $\geq 3$	6.1	-32	16

<sup>a</sup> From Simándi and Jáky.<sup>6</sup> Similar results were also obtained for maleic acid and maleate ion. <sup>b</sup> From Wiberg and Geer.<sup>20</sup> These authors also showed that several other unsaturated compounds gave similar activation parameters. <sup>c</sup> This work. <sup>d</sup> From Samándi and Jáky.<sup>23</sup>

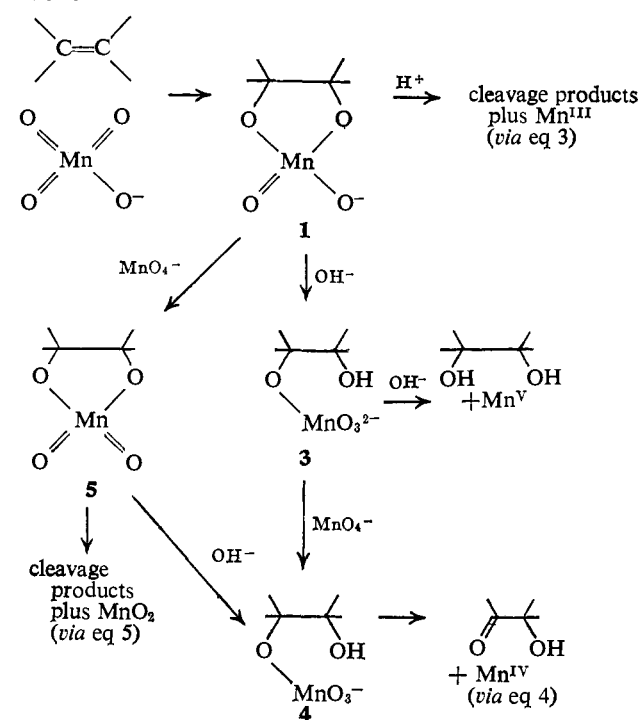
these reactions involve similar rate-determining steps, *i.e.*, formation of cyclic hypomanganate esters.

In summary it would appear that our results verify the suggestion first made nearly three-quarters of a century ago<sup>19</sup> that hypomanganate esters are formed as intermediates in the oxidation of alkenes by potassium permanganate. The transition state for the ester forming reaction appears to be one in which the olefinic carbons have acquired substantial  $sp^3$  character with the bond rearrangements nearly complete. The transition state for the second step (oxidative decomposition of the ester) probably contains a fairly well developed carbonyl group as evidenced by both isotope and substitution effects. Under these conditions the latter step is much slower and controls the overall rate of the reaction.

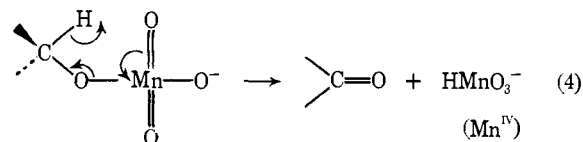
Accepting that hypomanganate esters are the initial intermediates in these reactions under all conditions, it is of interest to seek a mechanism which also satisfactorily accounts for the large array of products that can be obtained under various conditions. An attempt has been made to do this in Scheme I. The initial step under all conditions is apparently formation of the hypomanganate ester; however, the fate of the ester and consequently the products formed are determined almost entirely by the reaction conditions.

If the solution is acidic, the ester decomposes to give manganese(III), plus cleavage products. A possible reason why this reaction is observed only in acidic solutions may be found from a consideration of the stability of manganese(III). Under basic conditions it is not stable and consequently would be less easily formed there than in acidic solutions where it is known to be quite stable.<sup>18</sup>

**Scheme I**

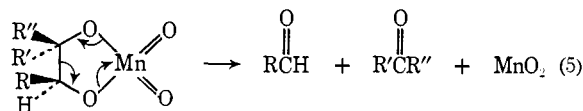


In the presence of base the ester would be hydrolyzed to 3 which could then undergo further hydrolysis yielding the diol or if the concentration of oxidant relative to base is high 3 may be oxidized to a manganate ester,<sup>11,12</sup> 4, which on decomposition would give the hydroxy ketone plus Mn<sup>IV</sup>. This latter step is reminiscent of the decomposition of chromium(VI) esters<sup>25</sup> and may be visualized as proceeding *via* eq 4. The competi-



tion between the step which leads to hydrolysis of 3 and the one which leads to oxidation explains why diols are obtained in best yields under conditions of high base but low permanganate concentration.<sup>2,3</sup> On the other hand, low base concentrations favor formation of the  $\alpha$ -hydroxy ketone.<sup>3,4</sup>

Alternatively the hypomanganate ester, 1, may be oxidized by permanganate or another cooxidant to the corresponding cyclic manganate ester, 5. It is likely that 5 would decompose to give cleavage products (eq 5) since it is known, for example, that manganate readily



cleaves diols.<sup>26</sup> It is also known that when a cooxidant such as periodate is present in relatively large concentrations double bonds are cleaved.<sup>27</sup> Although other explanations are available, the function of the cooxidant

(25) R. Stewart, "Oxidation Mechanisms," W. A. Benjamin, New York, N. Y., 1964, pp 33-57; H. Kwart and J. H. Nickle, *J. Amer. Chem. Soc.*, **95**, 3394 (1973).

(26) J. S. F. Pote and W. A. Waters, *J. Chem. Soc., London*, 717 (1956).

(27) R. U. Lemieux and E. von Rudloff, *Can. J. Chem.*, **33**, 1701, 1710 (1955); E. von Rudloff, *ibid.*, **33**, 1714 (1955); **34**, 1413 (1956).

may simply be that of converting the cyclic hypomanganate ester to the corresponding cyclic manganate ester.

It is not likely that the ester, **5**, would undergo decomposition to give hydroxy ketones as has been suggested elsewhere.<sup>8a</sup> The transfer of electrons that must occur during such a decomposition will proceed most readily if all of the orbitals involved are in the same plane. Since the C–C bond of **5** is in the same plane as the –O–Mn–O– bonds, cleavage could easily take place. On the other hand, cleavage of a C–H bond, which is required for formation of the  $\alpha$  ketol, would not likely

occur with the cyclic manganate ester because of the impossibility of the appropriate orbitals all becoming planar. Of course, if the ester were to be hydrolyzed before electron transfer took place the reaction depicted in eq 4 would probably take place.

**Acknowledgment.** The authors are pleased to acknowledge the financial assistance of the National Research Council of Canada in the form of an operating grant (D. G. L.) and a scholarship (J. R. B.). D. G. L. is also grateful for the hospitality of the Department of Chemistry at the University of Oslo where this paper was written.

## Intramolecular General Base Catalyzed Hydrolysis and Tertiary Amine Nucleophilic Attack *vs.* General Base Catalyzed Hydrolysis of Substituted Phenyl Quinoline-8- and -6-carboxylates

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**Abstract:** The reaction of a series of substituted (*p*-CH<sub>3</sub>O, H, *p*-Cl, *p*-CN, *o*-NO<sub>2</sub>, *p*-NO<sub>2</sub>, and 2,4-diNO<sub>2</sub>) phenyl quinoline-8- and -6-carboxylates (Q-8 and Q-6, respectively) with lyate species and tertiary amines has been investigated (20% (v/v) acetonitrile–H<sub>2</sub>O,  $\mu = 0.5$ , 30°). The Hammett  $\rho$  values for HO<sup>–</sup> attack on the Q-8 ( $\rho = 0.93$ ) and Q-6 ( $\rho = 0.88$ ) esters are of a magnitude anticipated for rate-determining nucleophilic addition to the ester group. The rate ratio  $k_{\text{HO}^- \text{Q-8}}/k_{\text{HO}^- \text{Q-6}} \simeq 0.2$  is attributed to the greater steric hindrance of the 8 isomers. Spontaneous hydrolysis ( $k_{\text{H}_2\text{O}}$ ) of the Q-8 esters greatly exceeds that for the Q-6 esters and exhibits a  $\rho$  value of 0.97. Employing two alternate means of approximation,  $k_{\text{H}_2\text{O} \text{Q-8}}/k_{\text{H}_2\text{O} \text{Q-6}} \cong 10^4$ . The enhanced rate of hydrolysis of the 8 isomers is attributed to intramolecular quinoline nitrogen general base assistance to nucleophilic attack of H<sub>2</sub>O in the case of the Q-8 esters. The reaction of tertiary amines with the Q-8 and Q-6 esters provides Hammett plots having  $\rho = 1.2$ – $1.5$  with the 2,4-diNO<sub>2</sub> substituted esters exhibiting marked positive deviations. From the similarity of the  $\rho$  values to those for HO<sup>–</sup> and H<sub>2</sub>O catalyzed hydrolysis, values of  $k_{\text{nH}_2\text{O}}/k_{\text{nD}_2\text{O}} = 1.5$ – $1.6$ , and  $\beta = 0.69$  for reaction of *p*-NO<sub>2</sub>-Q-6 with quinuclidine derivatives, tertiary amine catalysis of ester hydrolysis (except for the most reactive 2,4-diNO<sub>2</sub> substituted esters) is best ascribed to general base catalyzed attack of water. For tertiary amine catalyzed hydrolysis of the 2,4-diNO<sub>2</sub> substituted esters direct nucleophilic attack upon the esters is in effect: positive deviation from the  $\sigma\rho$  plot indicating a change in mechanism rather than a change in the rate-limiting step,  $\rho > 2.0$ ,  $k_{\text{nH}_2\text{O}}/k_{\text{nD}_2\text{O}} = 0.8$ – $0.9$ , and  $\beta = 0.95$  for reaction of quinuclidines with 2,4-diNO<sub>2</sub>-Q-6. Steric hindrance to approach of a tertiary amine to the 2,4-diNO<sub>2</sub> substituted esters (triethylamine and 4-morpholineethanesulfonic acid) brings about a change in the role of the tertiary amine from a nucleophilic to a general base catalyst. The rate ratios  $k_{\text{nQ-8}}/k_{\text{nQ-6}}$  range from 0.6 to 0.8 for the less reactive esters. For the 2,4-diNO<sub>2</sub> substituted esters the rate ratios increase with decreasing amine  $pK_a$  and range from 1.4 to 79. This observation has led to the suggestion that the transition state of the 2,4-diNO<sub>2</sub> substituted 8 isomer is electrostatically stabilized by the quinoline nitrogen lone pair. Reaction with imidazole is characterized by a linear Hammett plot with the 2,4-diNO<sub>2</sub> substituted esters, unlike with other tertiary amines, evidencing no deviation. The  $\rho$  values of 2.3 and 2.4 for imidazole catalyzed hydrolysis of the Q-8 and Q-6 esters, respectively, suggest that departure of phenoxide is at least partially rate limiting. The *p*-CH<sub>3</sub>O, H, *p*-Cl, and *p*-NO<sub>2</sub> substituted Q-6 esters are subject to imidazole catalysis of imidazole attack indicating that imidazole reacts directly with the esters as a nucleophile. This [imidazole]<sup>2</sup> rate term is seen only with the least reactive Q-8 ester (*p*-CH<sub>3</sub>O). The decreased importance of general base catalysis and the slight increase in bimolecular rate evidenced by the Q-8 esters in their reaction with imidazole suggest that the lone pair of the quinoline nitrogen may serve to stabilize the zwitterionic tetrahedral intermediate thereby increasing  $k_a$  sufficiently to make  $k_{\text{gb}}$  undetectable.

In a study of the hydrolysis of I, Fersht and Kirby found that the reaction proceeded *via* rate-limiting nucleophilic attack of water with intramolecular general base assistance from the carboxylate group.<sup>1</sup>

(1) A. R. Fersht and A. J. Kirby, *J. Amer. Chem. Soc.*, **89**, 4853 (1967).

Felton and Bruice<sup>2</sup> found that the hydrolysis of II and III occurred more rapidly than electronically equivalent isomers in which the nitrogen base is not located at the reaction site. Although this finding strongly suggested

(2) (a) T. C. Bruice and S. M. Felton, *J. Amer. Chem. Soc.*, **91**, 2799 (1969); (b) S. M. Felton and T. C. Bruice, *ibid.*, **91**, 6721 (1969).